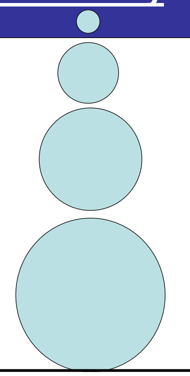
Periodic Trends

What pattern do you see?

-Number of fingers shown increases from left to right.

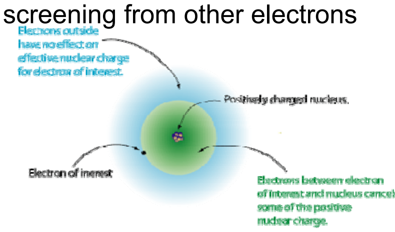
Elements in the Periodic Table are not arranged randomly

- Elements are arranged according to similar properties.
- As a result, there are several periodic trends that arise and can allow us to predict how elements will behave in certain circumstances.



Effective Nuclear Charge (Zeff)

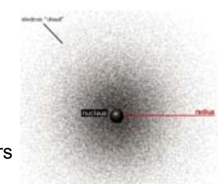
- is the **positive** charge that an electron experiences from the nucleus, equal to the nuclear charge but reduced by any shielding or screening from other electrons



1. ATOMIC RADIUS

•Atomic Radius: the estimate of the size of an atom from its nucleus to its outer perimeter.

- Atoms don't physically have a well defined boundary that we can conveniently present as a solid circle.
- Rather, this boundary is fuzzy → can't really measure atom size individually
- However, radius of an atom can be determined based on the distance between 2 atoms in compounds
- Atomic radii are measured in picometers (pm) 1pm = 10^{-12} m



ATOMIC RADIUS

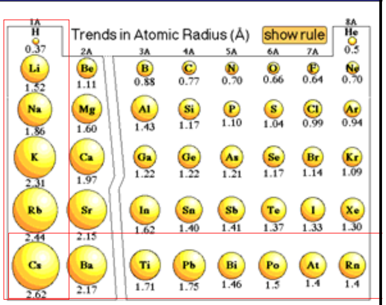
- What do you notice about the atomic radii as you move **across** a period?

•Atomic radii ↓

•move **down** a group?

•Atomic radii ↑

WHY?



Explaining Trends in Atomic Radius

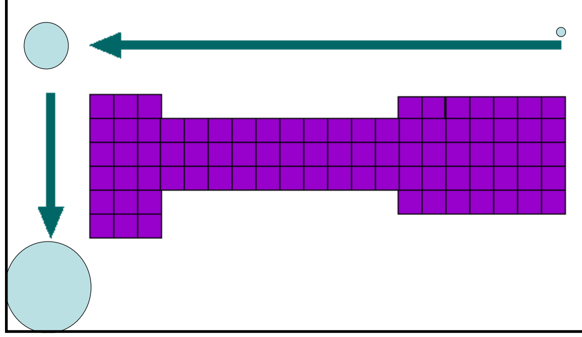
- From **left to right** within the same shell (or energy level) the atomic radius **decreases**:
- The number of **protons** increases by 1 as we move from one element to the next → effective nuclear charge **increases**
- # of **inner core electrons** remains **constant**
- **Valence** (or outermost) electrons are strongly attracted to the nucleus → **decrease** in size of atom

Explaining Trends in Atomic Radius

- **As you move down a group,** the atomic radius **increases**
- More electrons; thus more shells are added
- Outermost electrons feel a **repulsion by inner electrons** → effective nuclear charge **increases**
- Outermost electrons further away from nucleus → atom size **increases**
- Reduction in attractive force due to the inner electron is called the ***screening effect.***

Summary of Atomic Radius Trends

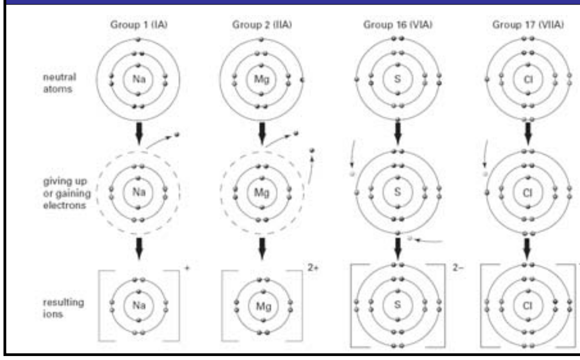
Atomic Size Increases With Arrows



Trends in Atomic Radius

Set of Atoms	Biggest Radius		**Smallest Radius**
a. Li, C, F			
b. Li, Na, K			
c. Ge, P, O			
d. C, N, Si			

Forming Ions (IONIZATION)



2. Ionization Energy (IE)

Ionization energy: The amount of energy required to remove **an electron from a gaseous atom**.

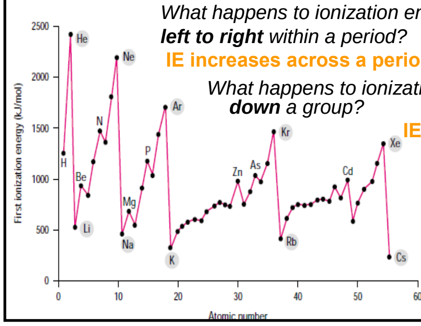
Ionization energy can be written as follows:

Na (g) + energy → Na+ + e- (the electron is now separated from Na)

The first ionization energy: The energy required to remove the least attracted electron from a gaseous atom of that element. This least attracted electron is in the outer shell of the atom.

Second ionization energy = amount of energy required to remove a second e- from a gaseous atom

Trends in Ionization Energy



*What happens to ionization energy (IE) from **left to right** within a period?*

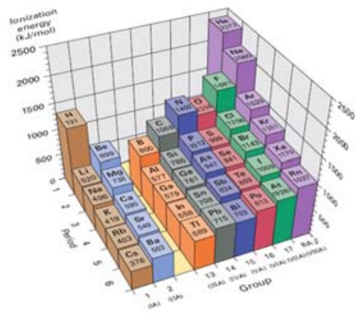
IE increases across a period

*What happens to ionization energy (IE) **down** a group?*

IE decreases

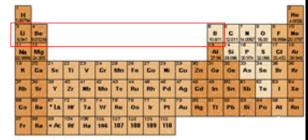
WHY?

First ionization energies for main group elements



Trends in Ionization Energy (across a period)

- As you move **across** a period, ionization energy **increases**
- More **electrons** being put in the same shell while # **protons** in nuclei increases
- As atom size decreases from left to right, outermost electrons experience **stronger** attractive forces from the nucleus
- Harder to remove them thus require more energy to remove these valence

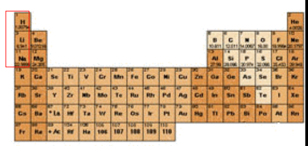


Trends in Ionization Energy (down a group)

- As you move **down** a group, ionization energy decreases
- Atom size **increases**
- Outermost electron become more distant from the nuclei → less nuclear force experienced
- Requires less **energy** to remove these e- → IE **decreases**
- **The lower the value, the more likely it is to lose an electron and become a positive ion.**

$H(g) + 131kJ \rightarrow H^+ + e^-$

Na(g) + 496kJ → Na+ + e- (the electron is now separated from Na)



Trends in Ionization Energy

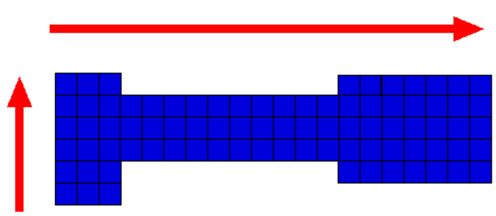
TABLE 1.9 Ionization Energies for Lithium, Beryllium, and Boron ($MJ \cdot mol^{-1}$)

Element	Ionization Energies				
	1st	2nd	3rd	4th	5th
lithium	0.5	7.3	11.8		
beryllium	0.9	1.8	14.8	21.0	
boron	0.8	2.4	3.7	25.0	32.8

- **What trend do you see?**
- **IE increases as successive e- are removed. Why?**
- After the first e- removed, an atom becomes an **ion** which is **positively** charged
- This indicates the **# protons** is greater than the **# electrons** in the ion
- Thus **greater nuclear charge** experienced by outer electrons making it **harder** (i.e. more energy required) to remove them from the ion.

Ionization Energy Summary: Opposite of Atomic Radius!

Ionization Energy Increases With Arrows

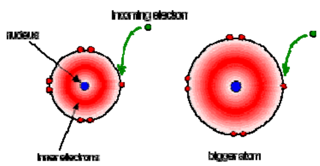


Trends in Ionization Energy

Set of Atoms	Most Ionization Energy		Least Ionization Energy
a. Mg, Si, S			
b. Mg, Ca, Ba			
c. F, Cl, Br			
d. Ba, Cu, Ne			
e. Si, P, N			

3. Electron affinity (EA)

- The energy **released** when an electron is **added** to a neutral, gaseous atom to form a anion (i.e. negatively charged ion)
- Ex: $F_{(g)}$ + e- → $F^{-}_{(g)}$ + energy (the electron has been added to the atom and energy was released)



Trends in Electron Affinity (EA)

- Trend observed?
- **EA increases from left to right and decreases down a group**
- Larger value of EA means it's **easier** to add e- to that atom.

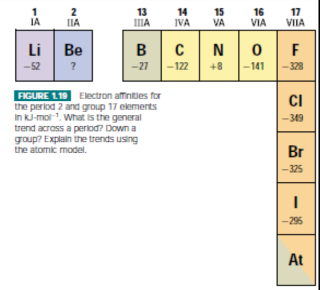
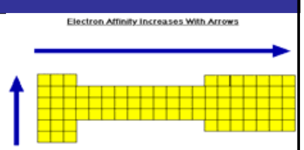


FIGURE 1.19 Electron affinities for the period 2 and group 17 elements in kJ·mol^{-1}. What is the general trend across a period? Down a group? Explain the trends using the atomic model.

Electron Affinity: The Why

- **A very negative value** of electron affinity indicates that a lot of energy is released when an electron is added to a gaseous atom, and that such a process is very likely to occur.



1 (IA)	2 (IIA)	13 (IIIA)	14 (IVA)	15 (VA)	16 (VIA)	17 (VIIA)	18 (VIIIA)
H −72.8							He (+21)
Li −59.6	Be (+241)	B −26.7	C −122	N 0	O −141	F −328	Ne (+29)
Na −52.9	Mg (+230)	Al −42.5	Si −134	P −72.0	S −200	Cl −349	Ar (+34)
K −48.4	Ca (+156)	Ga −28.9	Ge −119	As −78.2	Se −195	Br −325	Kr (+39)
Rb −46.9	Sr (+167)	In −28.9	Sn −107	Sb −103	Te −190	I −295	Xe (+40)
Cs −45.5	Ba (+52)	Tl −19.3	Pb −35.1	Bi −91.3	Po −183	At −270	Rn (+41)

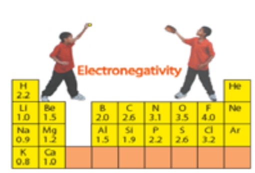
Figure 2.18 The units for electron affinity are the same as the units for ionization energy: kJ/mol. High negative numbers mean a high electron affinity. Low negative numbers and any positive numbers mean a low electron affinity.

Trends in Electron Affinity (EA)

Set of Atoms	Most Electron Affinity		Least Electron Affinity
a. Li, C, F			
b. C, O, Ne			
c. K, Si, O			
e. S, F, He			

4. ELECTRONEGATIVITY (EN)

- The ability of an atom to **attract** the shared **electrons** towards itself in a bond



The little black numbers indicate the electronegativity value for the element in the square.
A high number means a high electronegativity and therefore a stronger attraction for electrons.

Most electronegative atom is Fluorine whose value was assigned as 4.0 by Linus Pauling

TRENDS IN ELECTRONEGATIVITY

TABLE 1.10 Electronegativity Values for Main Group Elements

1 IA	2 IIA	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA
H 2.1						
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.9	Bi 1.9		

TRENDS IN ELECTRONEGATIVITY

Electronegativity increases from left to right. Why?

- From left to right, # protons **increases**. Shared electrons will be **attracted** more **strongly**

Electronegativity decreases down a group. Why?

- Down a group, # protons **increases** but also **more** inner electrons are added resulting in a shielding effect. Thus, shared electrons are **less** attracted to the nucleus.

TRENDS IN ELECTRONEGATIVITY

Set of Atoms	Most Electronegative		Least Electronegative
a. F, Cl, I			
b. Mg, Si, Cl			
c. K, Cl, Ne			

SUMMARY OF PERIODIC TRENDS